<u>Suggested Answers to Self–Check Questions</u>

1 **(a)** *Refer to IVY Solubility Equilibria Check Point 1 (Question 1) for worked solution.* K_{sp} = <u>1.77 x 10^{-14} mol^4 dm^{-12}</u>.

(b) *Refer to IVY Solubility Equilibria Check Point 1 (Question 2) for worked solution.* Solubility = <u>0.0729 g dm^{-3}</u>.

(c) Let the solubility of $Ca_3(PO_4)_2$ in water be s mol dm^{-3}.

$Ca_3(PO_4)_2(s) \rightleftharpoons 3Ca^{2+}(aq) + 2PO_4^{3-}(aq)$

At equilibrium in the saturated solution, $[Ca^{2+}]$ = 3s mol dm^{-3}

$[PO_4^{3-}]$ = 2s mol dm^{-3}

$$K_{sp} \text{ of } Ca_3(PO_4)_2 = [Ca^{2+}]^3[PO_4^{3-}]^2 = (3s)^3(2s)^2 = 1.0\times10^{-26} \text{ mol}^5 \text{ dm}^{-15}$$

$$108s^5 = 1.0\times10^{-26}$$

$$s = \left(\frac{1.0\times10^{-26}}{108}\right)^{1/5} = 2.474\times10^{-6} \text{ mol dm}^{-3}$$

Hence $[Ca^{2+}]$ in the saturated solution = (3)(2.474 x 10^{-6}) = <u>7.42 x 10^{-6} mol dm^{-3}</u>

2 *Refer to IVY Solubility Equilibria Check Point 2 for explanation.* Since <u>ionic product < K_{sp}</u>, precipitation of LiF will <u>not</u> occur.

3 **(a)** Upon mixing the two solutions and assuming no reaction,

$$[Ca^{2+}] = \left(\tfrac{2}{1000}\times0.50\right)\div\tfrac{2+1}{1000} = 0.3333 \text{ mol dm}^{-3} \qquad [NaOH] = [OH^-] = \left(\tfrac{1}{1000}\times2.0\right)\div\tfrac{2+1}{1000} = 0.6667 \text{ mol dm}^{-3}$$

$$\text{Ionic product} = \left[Ca^{2+}\right]\left[OH^-\right]^2 = (0.3333)(0.6667)^2 = \underline{0.148 \text{ mol}^3 \text{ dm}^{-9}}$$

Since <u>ionic product > K_{sp} of $Ca(OH)_2$</u>, precipitation of $Ca(OH)_2$ will occur.

A <u>white precipitate</u> will be observed in the test-tube.

(b) Upon mixing the two solutions and assuming no reaction,

$$[Ca^{2+}] = \left(\tfrac{2}{1000}\times0.50\right)\div\tfrac{2+1}{1000} = 0.3333 \text{ mol dm}^{-3} \qquad [NH_3] = \left(\tfrac{1}{1000}\times2.0\right)\div\tfrac{2+1}{1000} = 0.6667 \text{ mol dm}^{-3}$$

$NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$

At equilibrium, $[NH_4^+] = [OH^-]$

$[NH_3]_{eqm} \approx [NH_3]_{initial}$ = 0.6667 mol dm^{-3}, since NH_3 is a weak base with a small K_b,

$$K_b = \frac{\left[NH_4^+\right]\left[OH^-\right]}{\left[NH_3\right]} = \frac{\left[OH^-\right]^2}{\left[NH_3\right]}$$

$$\left[OH^-\right] = \sqrt{(1.74\times10^{-5})(0.6667)} = 3.406\times10^{-3} \text{ mol dm}^{-3}$$

$$\text{Ionic product} = \left[Ca^{2+}\right]\left[OH^-\right]^2 = (0.3333)(3.406\times10^{-3})^2 = \underline{3.87\times10^{-6} \text{ mol}^3 \text{ dm}^{-9}}$$

Since <u>ionic product < K_{sp} of $Ca(OH)_2$</u>, precipitation of $Ca(OH)_2$ will <u>not</u> occur.

<u>No white precipitate</u> will be observed in the test-tube.